

TOPIC 05 – ENERGETICS

5.1–5.2: HEAT CHANGE

IB Chemistry
T05D01



5.1 Exothermic and endothermic reactions - 1 hour

- 5.1.1 Define the terms exothermic reaction, endothermic reaction and standard enthalpy change of reaction ΔH_f^θ . (1)
- 5.1.2 State that combustion and neutralization are exothermic processes. (1)
- 5.1.3 Apply the relationship between temperature change, enthalpy change and the classification of a reaction as endothermic or exothermic. (2)
- 5.1.4 Deduce, from an enthalpy level diagram, the relative stabilities of reactants and products, and the sign of the enthalpy change for the reaction. (3)



5.2 Calculation of enthalpy changes - 3 hours

- 5.2.1 Calculate the heat energy change when the temperature of a pure substance is changed. (2)
- 5.2.2 Design suitable experimental procedures for measuring the heat energy changes of reactions. (3)
- 5.2.3 Calculate the enthalpy change for a reaction using experimental data on temperature changes, quantities of reactants and mass of water. (2)
- 5.2.4 Evaluate the results of experiments to determine enthalpy changes. (3)



5.3 Hess's law - 2 hours

- 5.3.1 Determine the enthalpy change of a reaction that is the sum of two or three reactions with known enthalpy changes. (3)



5.4 Bond enthalpies - 2 hours

- 5.4.1 Define the term average bond enthalpy. (1)
- 5.4.2 Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic. (3)



HL: 15.1 Standard enthalpy changes of reaction - 1.5 hour

- 15.1.1 Define and apply the terms standard state, standard enthalpy change of formation, ΔH_f^θ and standard enthalpy change of combustion ΔH_c^θ . (2)
- 15.1.2 Determine the enthalpy change of a reaction using standard enthalpy changes of formation and combustion. (3)



HL 15.2 Born–Haber cycle

- 2.5 hours

- 15.2.1 Define and apply the terms lattice enthalpy and electron affinity. (2)
- 15.2.2 Explain how the relative sizes and the charges of ions affect the lattice enthalpies of different ionic compounds. (3)
- 15.2.3 Construct a Born–Haber cycle for group 1 and 2 oxides and chlorides, and use it to calculate an enthalpy change. (3)
- 15.2.4 Discuss the difference between theoretical and experimental lattice enthalpy values of ionic compounds in terms of their covalent character. (3)



HL 15.3 Entropy - 1.5 hours

- 15.3.1 State and explain the factors that increase the entropy in a system. (3)
- 15.3.2 Predict whether the entropy change (ΔS) for a given reaction or process is positive or negative. (3)
- 15.3.3 Calculate the standard entropy change for a reaction (ΔS^θ) using standard entropy values (S^θ). (2)



HL 15.4 Spontaneity

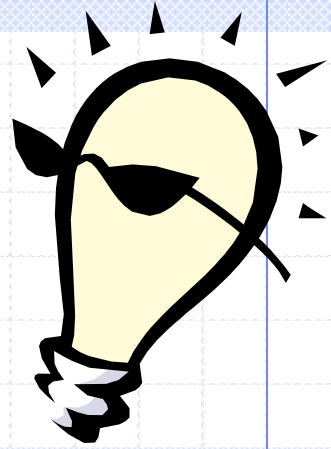
- 2.5 hours

- 15.4.1 Predict whether a reaction or process will be spontaneous by using the sign of ΔG^θ . (3)
- 15.4.2 Calculate ΔG^θ for a reaction using the equation $\Delta G^\theta = \Delta H^\theta - T\Delta S^\theta$ and by using values of the standard free energy change of formation, ΔG_f^θ . (2)
- 15.4.3 Predict the effect of a change in temperature on the spontaneity of a reaction using standard entropy and enthalpy changes and the equation $\Delta G^\theta = \Delta H^\theta - T\Delta S^\theta$. (3)

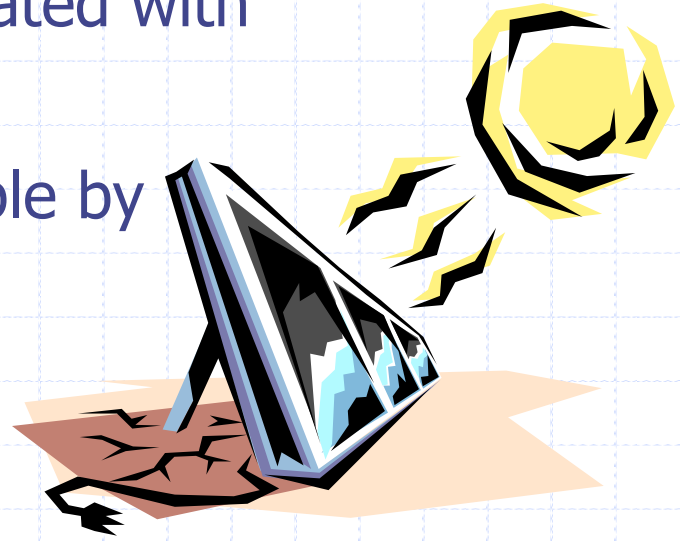


5.1 – Energy Transfer

Energy is the capacity to do work



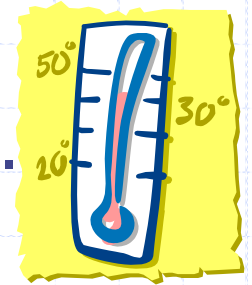
- ***Thermal energy*** is the energy associated with the random motion of atoms and molecules
- ***Chemical energy*** is the energy stored within the bonds of chemical substances
- ***Nuclear energy*** is the energy stored within the collection of neutrons and protons in the atom
- ***Electrical energy*** is the energy associated with the flow of electrons
- ***Potential energy*** is the energy available by virtue of an object's position



5.1 - Energy Changes in Chemical Reactions

Heat is the transfer of **thermal energy** between two bodies that are at different temperatures.

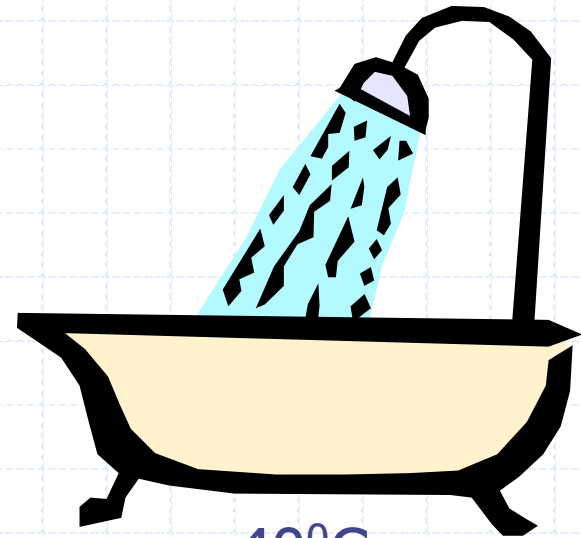
Temperature is a measure of the **thermal energy**.



Temperature \neq Thermal Energy



90°C



40°C

greater thermal energy
(since the mass is larger)



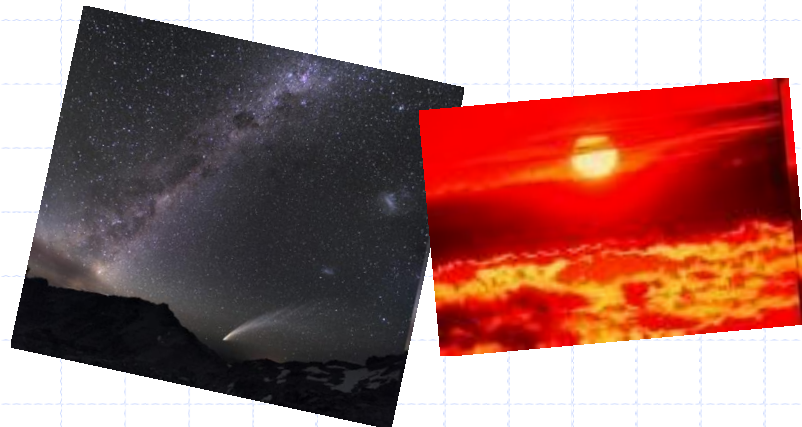
5.1 - Heat and Temperature

- ◆ **Heat** is energy that is transferred from one object to another due to a difference in temperature
- ◆ **Temperature** is a measure of the average kinetic energy of a body
- ◆ Heat is always transferred from objects at a higher temperature to those at a lower temperature



5.1 - Factors Affecting Heat Quantities

- ◆ The amount of heat contained by an object depends primarily on three factors:
 - The mass (m) of material
 - The temperature (t)
 - The kind of material and its ability to absorb or retain heat or (C).



5.1 - Heat Quantities

- ◆ The heat required to raise the temperature of 1.00 g of water 1 °C is known as a calorie.
- ◆ The SI unit for heat is the joule. It is based on the mechanical energy requirements.
- ◆ **1.00 calorie = 4.184 Joules**
- ◆ The energy required to raise 1 pound of water of 1 °F is called a **British Thermal Unit** or **BTU**
- ◆ The BTU is widely used in the USA to compute energy capacities of heating and air conditioning equipment



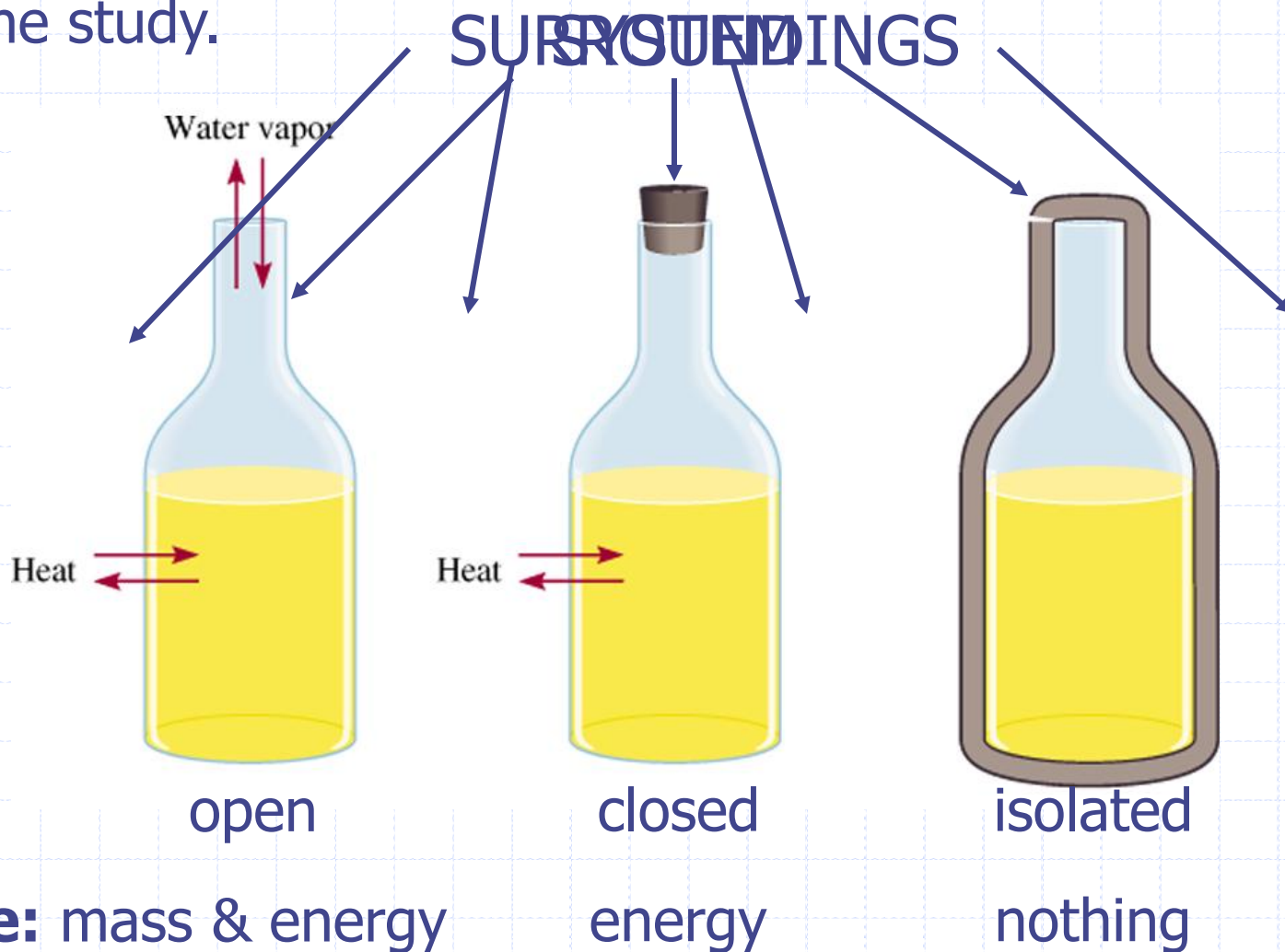
5.2 - Calorimetry

- ◆ Calorimetry involves the measurement of heat changes that occur in chemical processes or reactions.
- ◆ The heat change that occurs when a substance absorbs or releases energy is really a function of three quantities:
 - The mass
 - The temperature change
 - The heat capacity of the material



Thermochemistry is the study of heat change in chemical reactions.

The **system** is the specific part of the universe that is of interest in the study.

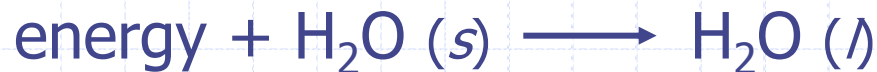


5.1 - Movement of Thermal Energy

Exothermic process is any process that gives off heat – transfers thermal energy from the system to the surroundings.



Endothermic process is any process in which heat has to be supplied to the system from the surroundings. Takes in heat.



5.2 - Equations for Heat Change (q)

The ***specific heat*** (c) of a substance is the amount of **heat** (q) required to raise the temperature of **one gram** of the substance by **one degree** Celsius.

Usually use specific heat, but you can use the ***heat capacity*** (C) as well [c substance is the amount of **heat** (q) required to raise the temperature of a given quantity in **mass** (m) of the substance by **one degree** Celsius.]

$$C = mc$$

Table 6.1 The Specific Heats of Some Common Substances

Substance	Specific heat (J/g · °C)
Al	0.900
Au	0.129
C (graphite)	0.720
C (diamond)	0.502
Cu	0.385
Fe	0.444
Hg	0.139
H ₂ O	4.184
C ₂ H ₅ OH (ethanol)	2.46

Heat (q) absorbed or released:

$$q = mc\Delta t$$

$$q = C\Delta t$$

$$\Delta t = t_{\text{final}} - t_{\text{initial}}$$



5.2 - Specific Heat values for Some Common Substances

Substance	$C_J \text{ g}^{-1} \text{ K}^{-1}$	$C \text{ J mol}^{-1} \text{ K}^{-1}$
Water (<u>liquid</u>)	4.184	75.327
Water (steam)	2.080	37.47
Water (ice)	2.050	38.09
Copper	0.385	24.47
Aluminum	0.897	24.2
Ethanol	2.44	112
Lead	0.127	26.4

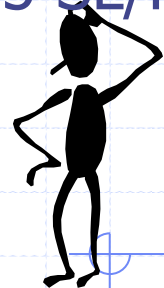


5.2 - Variables for Heat Change:

- ◆ **q**, is the heat change within the system
- ◆ **c**, is the specific heat of the substance
 - Does not take into account the mass
 - Use the equation $q = mc \Delta T$
- ◆ **C**, is the heat capacity of the substance
 - Takes into account the mass
 - Use the equation $q = C \Delta T$
- ◆ **m**, is the mass of the substance
- ◆ **ΔT** , is the change in temperature
 - $\Delta T = T_f - T_i$



5.2 - Heat Change Example:



Example: How much heat is given off when an 869 g iron bar cools from 94⁰C to 5⁰C? The specific heat of Fe is 0.444 J/g • ⁰C

$$c \text{ of Fe} = 0.444 \text{ J/g} \cdot ^\circ\text{C}$$

$$m \text{ of Fe} = 869 \text{ g}$$

$$\begin{aligned} \Delta t &= t_f - t_i \\ &= 5^\circ\text{C} - 94^\circ\text{C} \\ &= -89^\circ\text{C} \end{aligned}$$

$$\begin{aligned} q &= mc\Delta t \\ &= 869 \text{ g} \times 0.444 \text{ J/g} \cdot ^\circ\text{C} \times -89^\circ\text{C} \\ &= -34,000 \text{ J} \end{aligned}$$

We know the Fe cools (stated above) and the value given is -34 kJ, what does the negative value tell us? Is this an endo- or exo-thermic process?



5.2 - Calorimetry

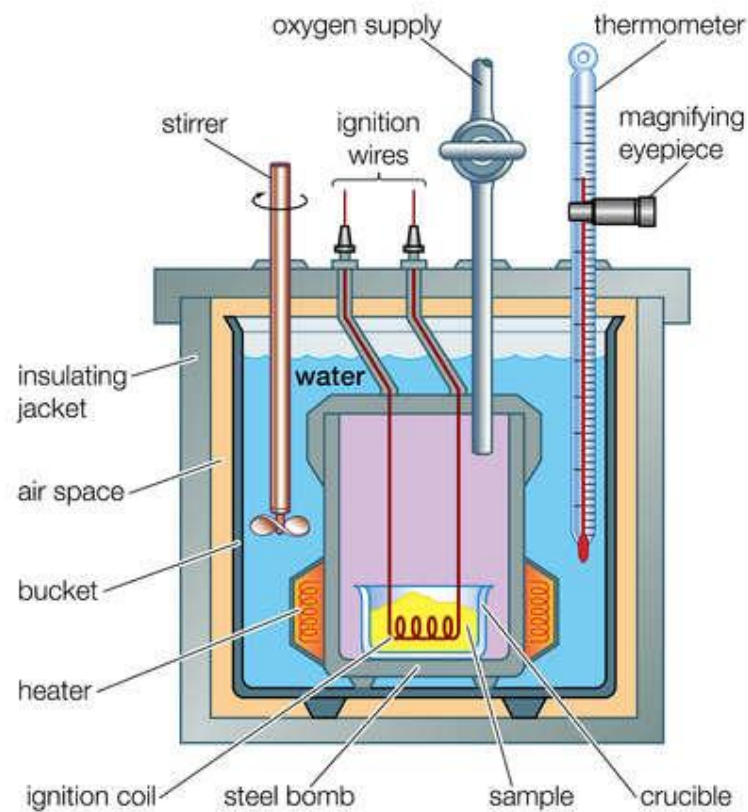
◆ Food is burned to completion after being ignited electrically (the bomb) and the heat released as it burns is measured by recording the temperature rise in a known mass of pure water.

◆ We use the equation

- $q = mc\Delta T$

- Energy lost = energy gained

◆ $q_{\text{food}} = - (q_{\text{H}_2\text{O}})$



Bomb Calorimeter

5.2 - Calorimetry Example

◆ A 0.78 g sample of a food substance was combusted in a bomb calorimeter and raised the temperature of 105.10 g of water from 15.4°C to 30.6°C. Calculate the energy value of the food in kJ/g.

- $m_{\text{H}_2\text{O}} = 105.10 \text{ g}$
- $\Delta T = T_f - T_i = 30.6 - 15.4 = 15.2^\circ\text{C}$ or 15.2 K
- $c_{\text{H}_2\text{O}} = 4.184 \text{ J/g}^\circ\text{C}$
- $q = mc\Delta T = (105.10 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(15.2^\circ\text{C})$
- $q = 6677.63 \text{ J}$ per 0.78 g sample heated
- Energy value = $6677.63 \text{ J} / 0.78 \text{ g}$
 - ◆ Energy = 8561.1 J/g OR 8.56 kJ/g



5.2 - Calorimetry Example 2

$$q = -(q)$$

- ◆ If 1.24 grams of NaOH was added to 50.00 grams of water and the temperature rose from 24.5°C to 44.5°C, provide the molar heat of solution for NaOH



5.2 - Energy lost = Energy Gained

◆ You needed to calculate q (heat change) for NaOH, but don't have all the information you need to solve directly, you must find heat change for water first:

- q_{NaOH} cannot be found directly, but $q_{\text{H}_2\text{O}}$ can be

- $q_{\text{NaOH}} = - (q_{\text{H}_2\text{O}})$ (this is the conservation of energy)

- $q_{\text{H}_2\text{O}} = mc\Delta T$

- ◆ $m = 50.00 \text{ g}$

- ◆ $s = 4.184 \text{ J/g}^\circ\text{C}$

- ◆ $\Delta T = 20.00 \text{ }^\circ\text{C}$

First Law of Thermodynamics!

Energy lost = Energy gained

$$q = -q$$

- $q_{\text{H}_2\text{O}} = (50.00 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(20.00^\circ\text{C})$

- $q_{\text{H}_2\text{O}} = 4,184 \text{ J}$

- $q_{\text{NaOH}} = -4,184 \text{ J}$



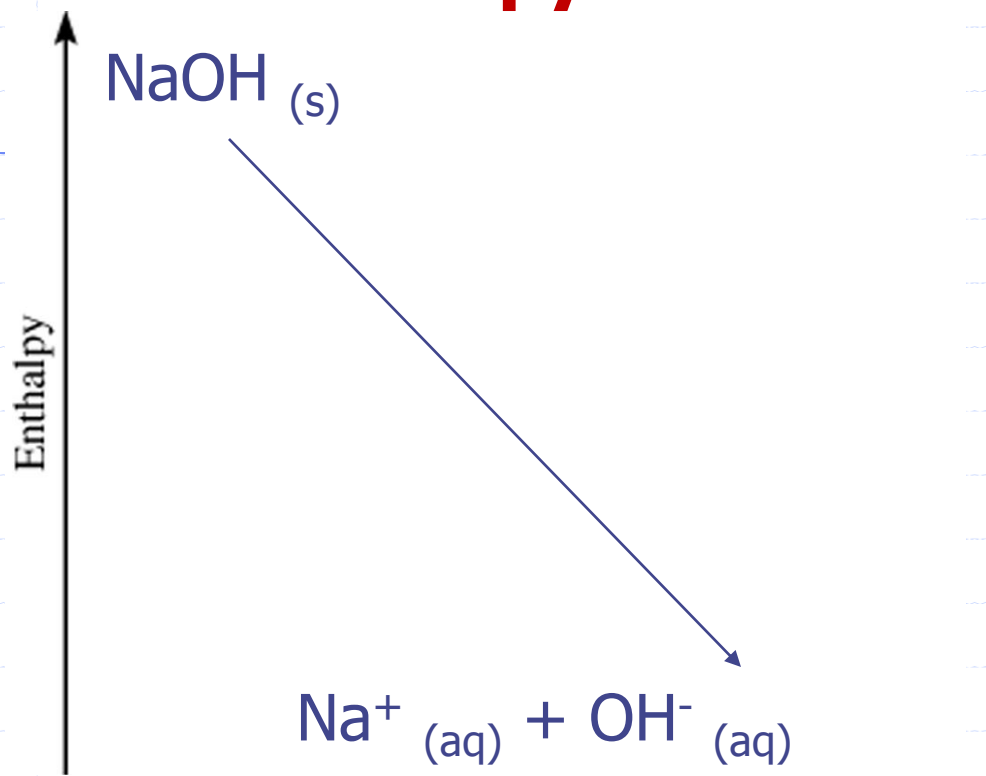
Now, how many grams of NaOH were used to produce this amount of heat? (1.24 g NaOH, next slide)

5.2 - Calculate Molar Heat of Solution

- ◆ 1.24 g NaOH produced a value of $q_{\text{NaOH}} = -4,184 \text{ J}$
- ◆ So, let's write this in units of Joules per gram
 - $q_{\text{NaOH}} = \underline{\hspace{2cm}}$
- ◆ You need to find Joules per mole (molar heats of solution) so that you can relate to the heat production of the other salt.
 - $\underline{\hspace{2cm}} \quad \underline{\hspace{2cm}} \quad \underline{\hspace{2cm}}$
 - $\text{NaOH(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \quad q_{\text{NaOH}} = -134 \text{ kJ/mol}$



5.1 - Enthalpy: Thermochemical Equations



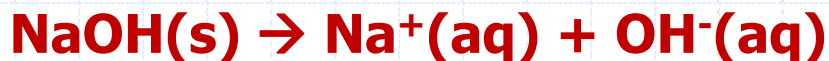
Is ΔH negative or positive?

System gives off heat

Exothermic

$$\Delta H < 0$$

134 kJ are released for every 1 mole of sodium hydroxide that is dissolved into water.



$$q_{\text{NaOH}} = -134 \text{ kJ/mol}$$
$$\Delta H = -134 \text{ kJ/mol}$$



5.1 - Enthalpy (H)

Enthalpy (H) is used to quantify the heat flow into or out of a system in a process that occurs at constant pressure. **Can be equated with q .** (q is generally in J and H in kJ)

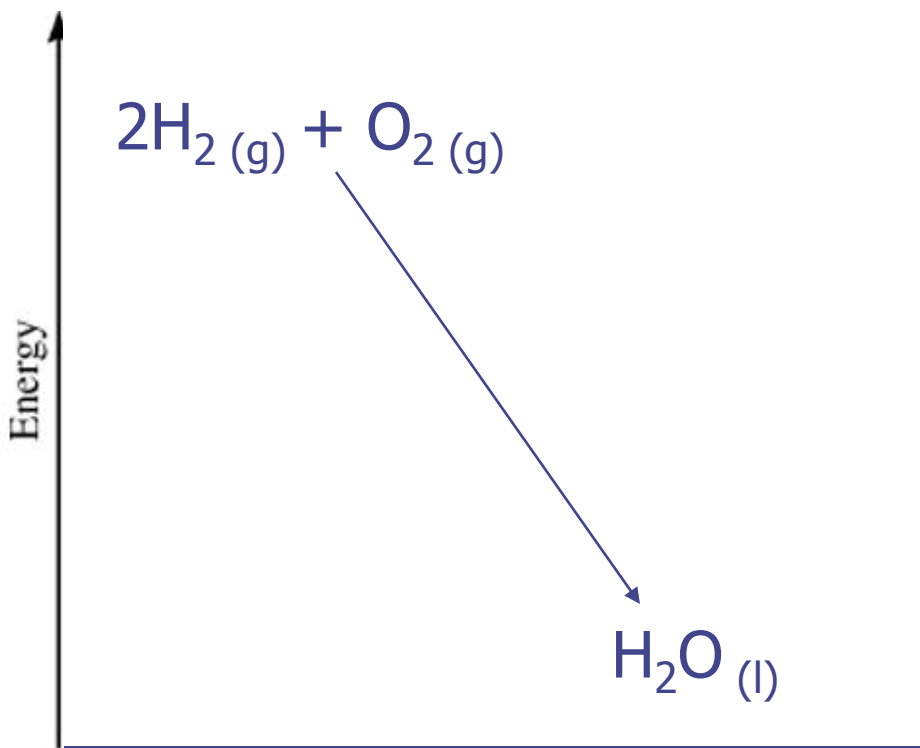
ΔH = heat given off or absorbed during a reaction **at constant pressure**

$$\Delta H = H(\text{products}) - H(\text{reactants})$$



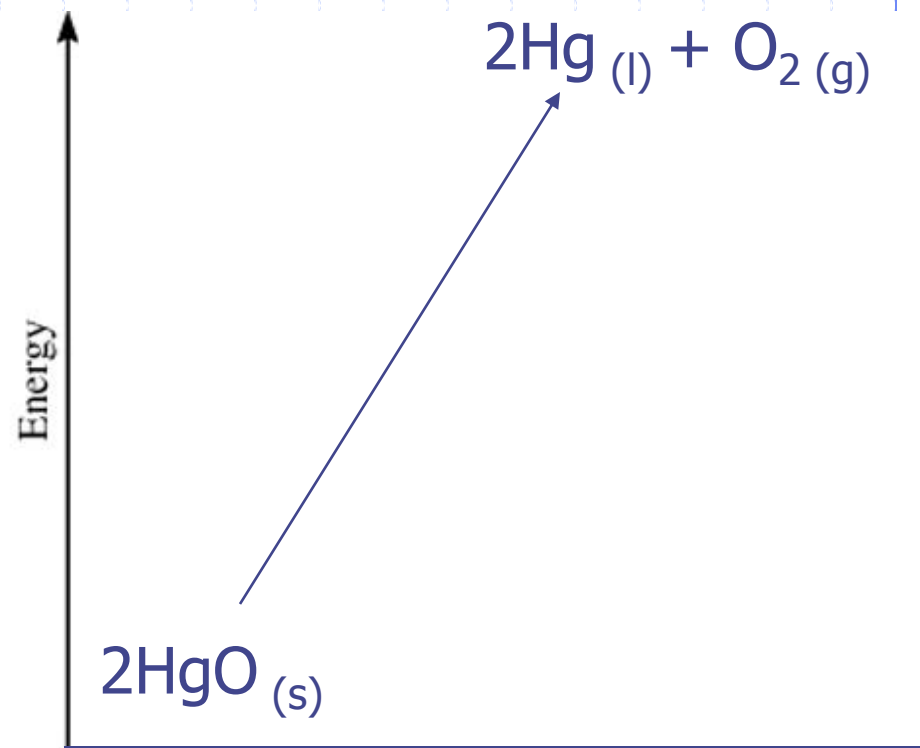
5.1 - Enthalpy (H)

$$\Delta H = H(\text{products}) - H(\text{reactants})$$



$$H_{\text{products}} < H_{\text{reactants}}$$

$\Delta H < 0$ (exothermic)

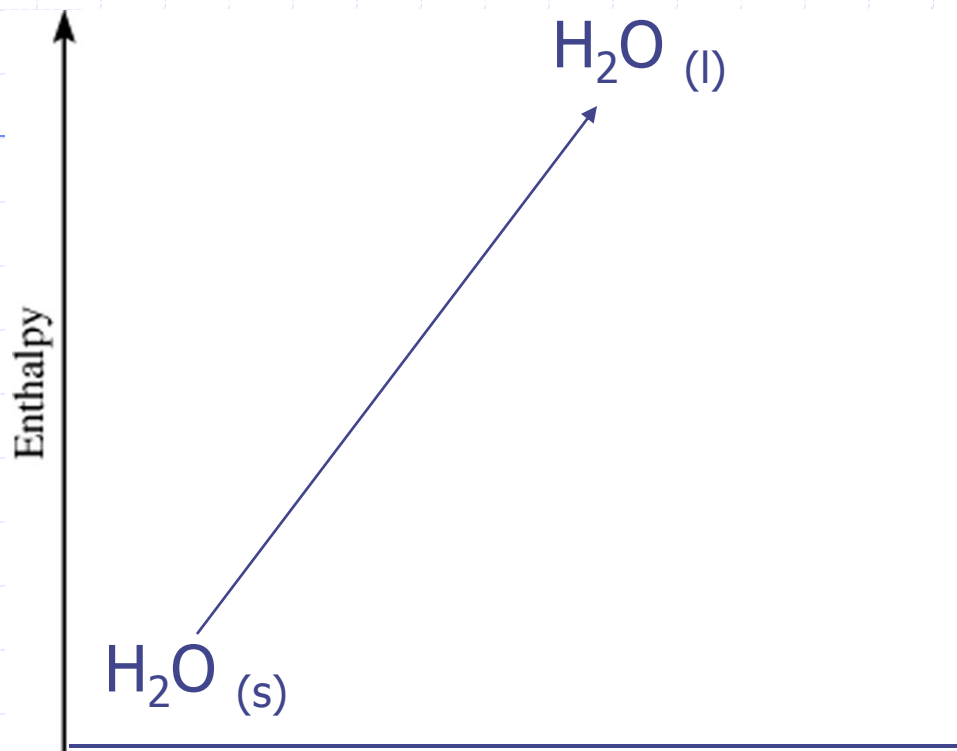


$$H_{\text{products}} > H_{\text{reactants}}$$

$\Delta H > 0$ (endothermic)



5.1 - Enthalpy: Thermochemical Equations



Is ΔH negative or positive?

System absorbs heat

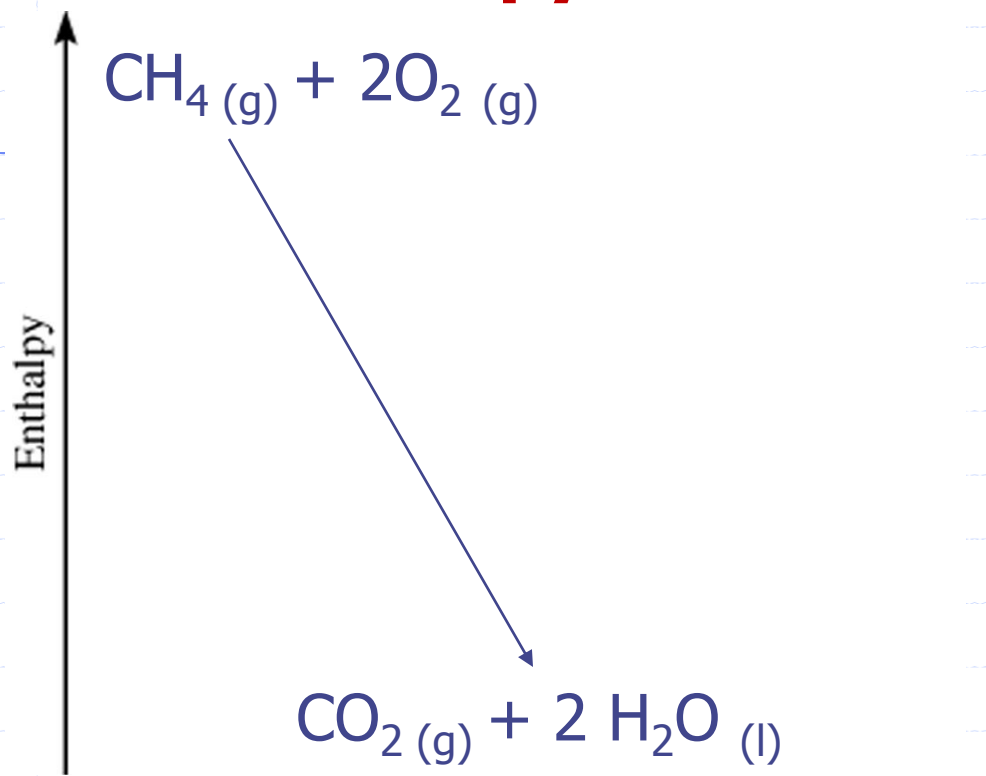
Endothermic

$$\Delta H > 0$$

6.01 kJ are absorbed for every 1 mole of ice that melts at 0°C and 1 atm.



5.1 - Enthalpy: Thermochemical Equations



Is ΔH negative or positive?

System gives off heat

Exothermic

$$\Delta H < 0$$

890.4 kJ are released for every 1 mole of methane that is combusted at 25°C and 1 atm.



5.3 - Enthalpy: Thermochemical Equations

- The stoichiometric coefficients always refer to the number of moles of a substance



- If you reverse a reaction, the sign of ΔH changes



- If you multiply both sides of the equation by a factor n , then ΔH must change by the same factor n .

